Chpt 15 Hmwk:
pg 462 Q1a, 3a, 4
468 Q1, 2
475 Q1, 2
477 Q11, 18, 19
478 Q22, 28, 30

06-02-06

Acids

"Chpt 15/16 Test starts ... here!"

- 5 major props of acids...
  1) contain H → release H^+ ions [in solution]
  2) strong electrolytes (produce ions in soln.)
  3) sour taste
  4) undergo neutralization rxns (w/ bases) ex. $HCl + NaOH \longrightarrow NaCl + H_2O$
  5) change the colors of indicator dyes
     - litmus { blue → red: acid; red → blue: base
     - Phenolphthalein { acidic → clear; basic → pink/red

• H^+ & H_3O^+ → interchangeable

Definitions of Acids (3):

[most specific] • Arrhenius Acid (AKA "the traditional acid") - any compound that contains H & produces H^+ (or H_3O^+) in solution.

ex. HCl, $HC_2H_3O_2$ • CH_4 → NOT an acid.

[more general] • Bronsted - Lowry Acid - any species (comp. or ion) that donates a H^+ or H_3O^+

(Arrhenius Acids are also Bronsted - Lowry Acids)

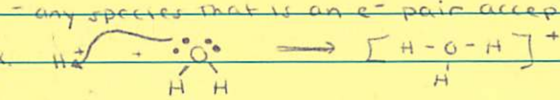
only works w/ Bronsted - Lowry Base (conjugate pairs)

ex. $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$ (H^+; the acid)

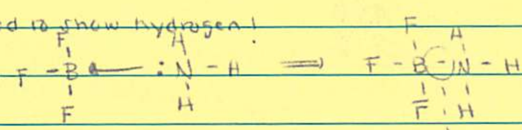
→ H^+ & H_3O^+ → b/c H_2O is added ⇒ same

ex. $HCO_3^- + H_2O \rightleftharpoons H_3O^+ + CO_3^{2-}$

[most general definition of an acid] • Lewis Acid - any species that is an e^- pair acceptor.

ex.


→ don't need to show hydrogen!

ex.


(violates the octet rule)

... coordinate covalent bond (any coordinate bond that there is only one donor for the e^- pair)

Strengths of Acids: strong vs. weak

- strong acids - ionize nearly completely ex. $HCl\,(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$

*** • 7 strong acids: HCl, $HClO_4$ (perchloric acid), $HClO_3$ (chloric acid), HBr (hydrobromic acid), HI (hydroiodic acid), H_2SO_4 (sulfuric acid), & HNO_3 (nitric acid)

[diprotic → 2 H's → gives off 2 protons]

- weak acids - acids do not ionize completely. ex. $HC_2H_3O_2\,(aq) \xrightleftharpoons{5\%} H^+(aq) + C_2H_3O_2^-(aq)$

• if not one of the strong acids... it's one of the weak acids. (common sense ☺)

↳ only ≈ 5% in ionization.

ex. H_2CO_3 (diprotic acid → protons are lost → ONE at a time)

$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$

06-05-06

- polyprotic acids → usually a weak acid (except for H_2SO_4)

$$\underbrace{H_2O + \underbrace{H_2CO_3}_{\text{weak acid}} \rightleftharpoons \underbrace{HCO_3^-}_{\text{strong base.}}}_{} + H_3O^+$$

conjugate pairs

* protons are lost one at a time!

Bases

· properties:

1) bitter taste
2) (like acids) are strong electrolytes
   ex $NaOH \rightarrow Na^+ + OH^-$
3) will change indicators
4) (like acids) will react with acids (bases for acids) ⟹ neutralization
5) dilute bases are slippery ex. soap

Definitions of Bases (3) → there is overlap in the definitions

[most specific]

1) **Arrhenius base** ("traditional" definition of a base) any compound that produces OH^- ions in aqueous solns.

   ex $NaOH(s) \longrightarrow Na^+(aq) + OH^-(aq)$

   $NH_3\uparrow + H_2O \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

2) **Brønsted-Lowry** - any specie that accepts a $\boxed{H^+}$

   ex. $NH_3\uparrow + H_2O \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

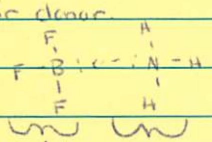
- **amphoteric** - a species that behaves as either an acid or a base. ex H_2O

  ex $\underbrace{HCO_3^-}_{\text{also amphoteric}} + H_2O \rightleftharpoons H_2CO_3 + OH^-$

[broadest definition]

3) **Lewis** - an e^- pair donor.

   ex.



acid [acceptor] base [donor]

☆ Brønsted-Lowry bases ACCEPT a proton & Lewis bases DONATE an e^- pair ☆

Strengths of Bases

- **strong base** - dissociate nearly completely ex $NaOH(s) \xrightarrow{} Na^+(aq) + OH^-(aq)$ ⟹ 100% dissociation

*** • Group 1 & 2 hydroxides are strong bases --> alkaline earth metals.

  - $Ca(OH)_2$ → a strong base, but it's insoluble... ?!
    - → trick → will dissolve → just a very low amount
    - reaches equilibrium at very low levels
    - but when does dissolve → dissolves fully.

- many will vary in solubility but still will nearly 100% ionize.

- **weak bases** - do not completely ionize. ex. $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

  $Fe(OH)_3(aq) \rightleftharpoons Fe(OH)_2^+ + OH^-$

Brønsted-Lowry Conjugate Acid-Base Pairs

ex. $\underbrace{HBr}_{\text{strong acid}} + \underbrace{H_2O}_{\text{base}} \xrightarrow{\text{b/c it's a strong acid}} \underbrace{H_3O^+}_{\text{acid}} + \underbrace{Br^-}_{\text{base}}$

6/6/06 Bronsted Lowry Conjugate Acid Base Pairs

Strong acid $\boxed{HBr}$ + H_2O $\longrightarrow$ $\boxed{Br^-}$ + H_3O^+

A B B A

Same anion / molecule but acid has 1 more H^+

$H_2SO_3 + H_2O \rightleftharpoons HSO_3^- + H_3O^+$

Weak acid A B B A

For B/L A/B Pairs - ① THE WEAKER SPECIES IS ALWAYS FAVORED

exception H_2SO_4

② Strong acid has weak conjugate base

$Br^- + H_2O \rightleftharpoons HBr + OH^-$

Rxn will most likely not happen

Water won't give up proton to weak base

Hydrolysis

$HSO_3^- + H_2O \rightleftharpoons H_2SO_3 + OH^-$

6/7/06 Reactions w/ Acids + Bases

① Neutralization Rxn Acid + Base → H_2O + Salt

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

$H^+ + Cl^- + Na^+ + OH^- \rightarrow H_2O + Na^+ + Cl^-$

Ionic equation (shows all ions present in chem Rxn)

Net Ionic Eqn

$H^+ + OH^- \rightarrow H_2O$

② Acids React w/ Metals → H_2 gas + Salt

$2HCl_{(aq)} + 2Li_{(s)} \rightarrow H_2\uparrow + 2LiCl$

③ Nonmetal Oxides React w/ H_2O → Acid

$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$

④ Acids React w/ Carbonates → $CO_2 + H_2O$ + Salt

(5 $HCl_{(aq)} + NiCO_3 \rightarrow CO_2\uparrow + H_2O + NiCl_2$

⑤ Acids React w/ metal oxides → H_2O + Salt

$HCl_{(aq)} + MgO_{(g)} \rightarrow H_2O_{(l)} + MgCl_{2(aq)}$

Chpt 16 Titration

<u>Titration</u> - a measurement of unknown solution concentration using known solution concentration

<u>solution of known [ ]</u>
standard solution

<u>Equivalent</u> - for an acid - proton
for a base OH^-

10.0 mL of NaOH of [unknown] <u>neutralized</u> are added to 15 mL of 0.50 M HCl. What is the [NaOH]?

$NaOH + HCl \rightarrow NaCl + H_2O$

0.015 L × 0.50 M HCl =
0.0075 mol HCl = 0.0075 mol H^+
0.0075 mol OH^-
↓
0.0075 mol NaOH

pH 7 $[H^+] = [OH^-]$
$H_2O \rightleftharpoons H^+ + OH^-$

higher pH
$[OH^-] > [H^+]$

lower pH
$[H^+] > [OH^-]$

6/13 pH

autoionization of H_2O

$H_2O \rightleftharpoons H^+ + OH^-$

in pure water @ 25°C

$[H^+] = 1.0 \times 10^{-7}\ M$
$[OH^-] = 1.0 \times 10^{-7}\ M$

Because of LeChattlier's Principle
inverse prop b/w $[H^+]$ and $[OH^-]$
$[H^+] \uparrow,\ [OH^-] \downarrow$

$[H^+] \times [OH^-] = k_w$
autoionization const. of H_2O

$k_w = 1.0 \times 10^{-14}\ M^2$

What is the $[H^+] + [OH^-]$ in a 0.010 M HCl?

$HCl \longrightarrow H^+ + Cl^-$

$[H^+] = 0.010\ M\ HCl$

$[OH^-] = \frac{k_w}{[H^+]}$

$\frac{k_w = 1.0 \times 10^{-7}\ M^{\cancel{2}}}{[H^+] = 0.010\ \cancel{M}}$

$= 1.0 \times 10^{-12}\ M\ OH^-$

Weak Acid
$HC_2H_3O_2$ (5% ionization)

Do problem normally then take 5% of answer

pH = potential of Hydrogen
↳ makes [H^+] and [OH^-] easier to handle

$pH = -\log[H^+]$

$-\log(.01\,M)$
$= 2$

Basic solution pH

$0.017\ M\ NaOH$

$[OH^-] = 0.017\ M$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}\ M^{\cancel{2}}}{0.017\ \cancel{M}} = 5.9 \times 10^{-13}\ M$

$pH = -\log(5.9 \times 10^{-13}\ M)$
$= 12.2$

$-\log\ (K_w) = ([H^+][OH^-])\ -\log$

$-\log K_w = -\log[H^+][OH^-]$
$-\log K_w = -\log(H^+) + -\log(OH^-)$
$-\log(1.0 \times 10^{-14}) = -\log(H^+) + -\log(OH^-)$
$14 = pH + pOH$

$14 = pH + 1.8$
$12.2 = pH$

6/14 Flag Day

If 13.0 mL of 3.8 M HCl titrates 15.0 mL of KOH to equivalence, what is the ~~K~~ [KOH]OH ?

$HCl + KOH \rightarrow$

mol H^+ = mol OH^-

$M_a V_a = M_b V_b$

$\frac{mol}{L} \times L = mol$

M_a = 3.8 M
V_a = 13.0 mL
M_b = ?
V_b = 15.0 mL

$$\frac{3.8M \cdot 13.0\ mL}{15.0\ mL} = 3.3\ M\ KOH$$

When 27.3 mL of 0.25 M H_2SO_4 neutralize 2.10 M NaOH, what volume of base was used?

$n_a = 2$
$n_b = 1$

$M_a = 0.25\ M\ H_2SO_4$
$V_a = 27.3\ mL$
$M_b = 2.10\ M\ NaOH$
$V_b = ?$

$H_2SO_4 + 2NaOH \rightarrow 2H_2O + Na_2SO_4$

$H_2SO_4 + OH^- \rightarrow H_2O + HSO_4^- + OH^-$
$\rightarrow 2H_2O + SO_4^{2-}$

mol H^+ ↓ ← mol OH^-

$n_a M_a V_a = n_b M_b V_b$

$$V_b = \frac{2 \cdot 0.25M \cdot 27.3\ mL}{1 \cdot 2.10 M} = 6.5\ mL$$